MOLAR VOLUME OF A GAS

Objective:

Determine the volume of 1 mole of hydrogen gas at STP using experimental data, known mathematical relationships, and a balanced chemical equation.

Pre-lab Discussion Notes:

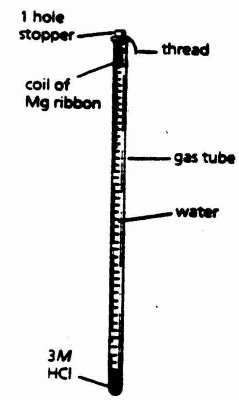

Figure 21-1

Materials:

gas-measuring tube
1000ml beaker
one-hole stopper (for gas tube)
10ml graduated cylinder
ring stand
utility clamp
thermometer
Mg ribbon
copper wire
metric ruler
safety glasses
400ml beaker
3 *M* hydrochloric acid

Procedure:

1. Obtain a piece of magnesium ribbon from your teacher. Clean it with steel wool and then measure its mass and record in the data table (a).
2. Obtain a piece of copper wire about 15 cm long. Wrap one end around the piece of magnesium ribbon, leaving about 10 cm of wire free. Bend the piece of magnesium so that it will fit easily into the gas-measuring tube.
3. Obtain about 10ml of 3 M hydrochloric acid (HCl). CAUTION: Handle acid with care! Carefully pour the HCl into a gas-measuring tube.
4. Tilt the gas-measuring tube slightly. Using a beaker, slowly fill the gas-measuring tube with water at room temperature. Try to avoid mixing the acid and water as much as possible.

5. Lower the piece of magnesium ribbon 4 or 5 cm into the gas-measuring tube. Bend the wire over the edge of the tube and insert the one-hole rubber stopper into the tube as shown in figure 21-1.
6. Add about 300ml of water at room temperature to a 400ml beaker. Set up a ring stand and utility clamp, and place the beaker of water in the position shown in Figure 21-2.

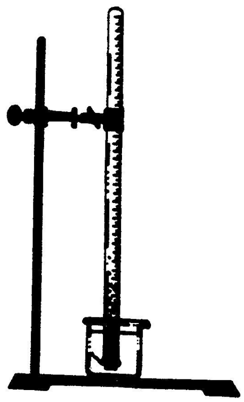

Figure 21-2

7. Place your finger over the hole in the rubber stopper and invert the gas-measuring tube. Lower the stoppered end of the tube into the beaker of water. Clamp the tube in place so that the stoppered end is a few centimeters above the bottom of the beaker (figure 21-2). Record your visual observations as (f) in the data table.
8. Let the apparatus stand about 5 minutes after the magnesium has completely reacted. Then, tap the sides of the gas-measuring tube to dislodge any gas bubbles that may have become attached to the sides of the tube. Place your finger over the hole in the stopper and transfer the tube to a 1000ml beaker filled with water. Lower the end of the tube into the water and remove your finger from the hole.
9. Move the tube up or down (to equalize pressure) until the water level in the tube is the same as that in the 1000ml beaker. On the scale of the gas-measuring tube, read the volume of the gases in the tube. Record this volume as (b) in your data table.
10. In the data table, record the room temperature, (c), and the barometric pressure, (d).
11. If time permits and your teacher indicates, repeat the experiment.

Observations and Data

	Trial 1	Trial 2	Trial 3
(a)Mass of Mg ribbon	g	g	g
(b)Volume of H_2 gas in tube	ml	ml	ml
(c)Room temperature	°C	°C	°C
(d)Barometric pressure	mm Hg	mm Hg	mm Hg
(e)Water vapor pressure at room temperature, P_{H2O}	mm Hg	mm Hg	mm Hg
(f)Visual observations:			

Calculations:

1. Calculate the number of moles of Mg reacted (which is equal to the number of moles of H_2 gas produced):

$$\text{Number of moles of Mg} = \frac{\text{mass of Mg used}}{\text{24.3g/mol (molar mass of Mg)}}$$

2. Find the pressure exerted by the H_2 gas in the tube:

$$P_{H2} = P_{barometric} - P_{H2O}$$

3. Convert room temperature from °C to Kelvin: K = °C + 273

4. Now, the volume of the H_2 gas you measured (b in your data table) was at room conditions – NOT standard temperature and pressure conditions. You are going to calculate what that volume would be if you moved your gas to a place of standard temperature and pressure (STP = 273K and 760mm Hg). To find the volume of the H_2 gas at STP use the combined gas law, which is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \textit{BUT WE NEED TO SOLVE FOR } V_2\textit{, SO:} \quad V_2 = \frac{(V_1P_1T_2)}{(T_1P_2)}$$

$P_1 = P_{H2}$ (from calculation 2)
V_1 = experimental value of H_2 (from b in your data chart)
T_1 = room temperature (K)
T_2 = 273 K
P_2 = 760mm Hg
V_2 = volume of H_2 at STP

5. You just found the volume of your gas at STP. Now, you want to find out what that volume would be if you had exactly 1 mole of it.

So, $\dfrac{\text{volume of } H_2 \text{ gas } (V_2)}{\text{moles of } H_2 \text{ gas (from calculation 1)}}$ will be set to: $\dfrac{x \text{ ml}}{1 \text{ mol}}$

solve for x = __________ml/mole and then convert to: ___________L/mole

Conclusions and Questions

1. The accepted value for the molar volume of a gas is 22.4 liters (22,400 ml). How does your experimentally determined value compare with this accepted value? Calculate your percentage error. Which trial had the least error?

2. What are some sources of error in this experiment? Be specific!

3. How many liters would the following number of moles of any gas occupy at STP?
a. 0.25 moles b. 0.5 mole c. 1 mole d. 2 moles e. 2.5 moles

4. Find the volume of the following masses of gases at STP:
a. 80g O_2 b. 10g H_2 c. 14g N_2 d. 66g CO_2

5. What happens to the other product of the reaction used in this experiment?

6. The following volumes of gas were measured at room temperature (293K) and at a pressure of 740mm Hg. What would these volumes be at STP?

   a. 32ml

   b. 7.1 L

   c. 56ml

   d. 35.5L

7. You measure the volume of 0.002 moles of hydrogen at STP to be 42ml. What would be the measured volume of 1 mole of the gas be?

8. What would be your percent error in #7? (Remember, the accepted value for 1 mole of gas at STP is 22.4L).